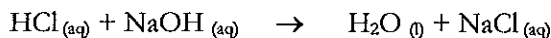


Acid-Base Equilibrium: Strong Acid – Strong Base Titrations

1. Consider the titration of 25.0 mL of 0.100M HCl with a 0.100 M NaOH solution.



Determine the pH at the following points:

- Initial pH (no NaOH added)
- pH after the addition of 10.00 mL of NaOH
- pH after the addition of 25.00 mL of NaOH
- pH after the addition of 35.00 mL of NaOH

$$V_e = \frac{C_a V_a}{C_b} = \frac{(0.100\text{M})(25.0\text{mL})}{0.100\text{M}} = \underline{25.0\text{mL}}$$

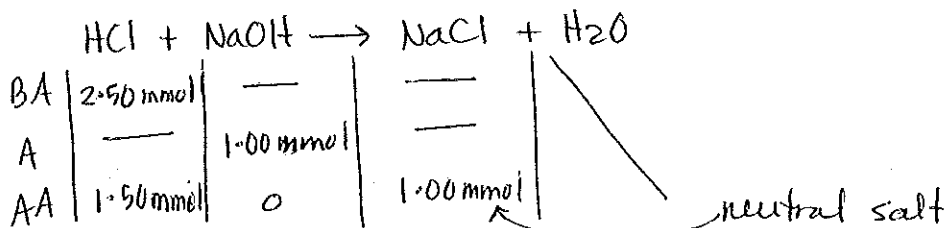
↑
equivalence volume

(a) strong acid solution

$$[\text{H}_3\text{O}^+] = [\text{HCl}] = 0.100\text{M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log(0.100) = \underline{1.000}$$

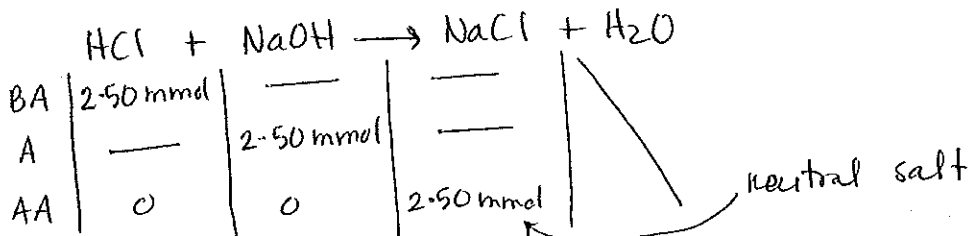
(b) Excess HCl:
(10.00 mL NaOH)



$$[\text{H}_3\text{O}^+] = [\text{HCl}] = \frac{1.50\text{ mmol}}{35.0\text{ mL}} = 0.0429\text{ M}$$

$$\text{pH} = -\log(0.0429) = \underline{1.368}$$

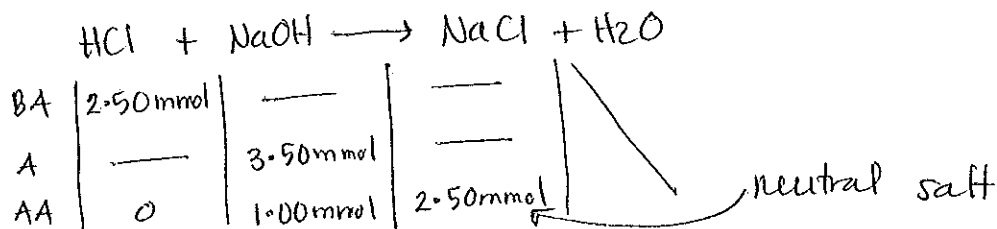
(c) Equivalence Point:
(25.00 mL NaOH)



$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7}\text{ M}$$

$$\text{pH} = -\log(1.0 \times 10^{-7}) = \underline{7.00}$$

(d) Excess NaOH:
(35.00 mL NaOH)



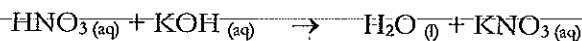
$$[\text{OH}^-] = [\text{NaOH}]_{\text{excess}} = \frac{1.00\text{ mmol}}{60.0\text{ mL}} = 0.0167\text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log(0.0167) = 1.778$$

$$\text{pH} = 14.00 - 1.778 = \underline{12.22}$$

Acid-Base Equilibrium: Strong Acid – Strong Base Titrations (Practice)

1. Consider the titration of 30.0 mL of 0.100M KOH with a 0.200 M HNO₃ solution.



Determine the pH at the following points:

$$V_e = \frac{C_b V_b}{C_a} = \frac{(30.0 \text{ mL})(0.100 \text{ M})}{0.200 \text{ M}} = \underline{15.0 \text{ mL}}$$

↑
equivalence volume

- Initial pH (no HNO₃ added)
- pH after the addition of 5.00 mL of HNO₃
- pH after the addition of 15.00 mL of HNO₃
- pH after the addition of 25.00 mL of HNO₃


(a) Strong base solution

$$[\text{OH}^-] = [\text{KOH}] = 0.100 \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log (0.100) = 1.000$$

$$\text{pH} = 14.00 - 1.000 = \underline{13.00}$$

(b) Excess KOH : HNO₃ + KOH → KNO₃ + H₂O
(5.00 mL HNO₃)

BA	_____	3.00 mmol	_____	
A	1.00 mmol	_____	_____	
AA	0	2.00 mmol	1.00 mmol	

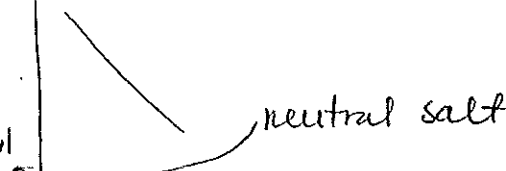
$$[\text{OH}^-] = \frac{2.00 \text{ mmol}}{35.0 \text{ mL}} = 0.0571 \text{ M}$$

$$\text{pOH} = -\log (0.0571) = 1.243$$

$$\text{pH} = 14.00 - 1.243 = \underline{12.76}$$

(c) Equivalence Point :
(15.00 mL HNO₃)




BA	_____	3.00 mmol	_____	
A	3.00 mmol	_____	_____	
AA	0	0	3.00 mmol	

$$[\text{H}^+] = [\text{H}_3\text{O}^+] = 1.0 \times 10^{-7} \text{ M}$$

$$\text{pH} = -\log (1.0 \times 10^{-7}) = \underline{7.00}$$

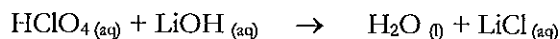
(d) Excess HNO₃ :
(25.00 mL HNO₃)

BA	_____	3.00 mmol	_____	
A	5.00 mmol	_____	_____	
AA	2.00 mmol	0	3.00 mmol	

$$[\text{H}^+] = [\text{HNO}_3]_{\text{excess}} = \frac{5.00 \text{ mmol}}{55.0 \text{ mL}} = 0.0909 \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (0.0909) = \underline{1.041}$$

2. Consider the titration of 20.0 mL of 0.150M HClO_4 with a 0.120 M LiOH solution.



Determine the pH at the following points:

- Initial pH (no LiOH added)
- pH after the addition of 10.00 mL of LiOH
- pH after the addition of 25.00 mL of LiOH
- pH after the addition of 35.00 mL of LiOH

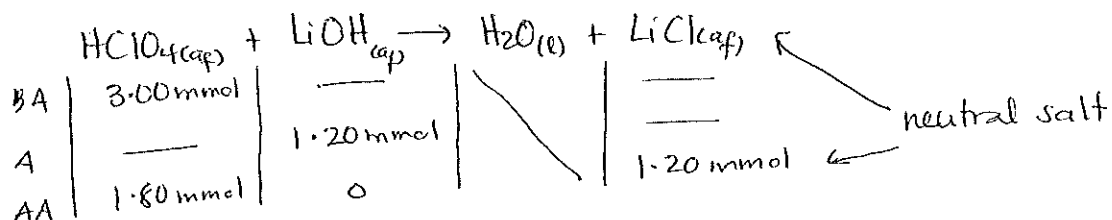
$$V_e = \frac{C_a V_a}{C_b} = \frac{(0.150M)(20.0\text{mL})}{0.120M} = \underline{25.0\text{mL}}$$

↑
equivalence volume

(a) Strong acid solution

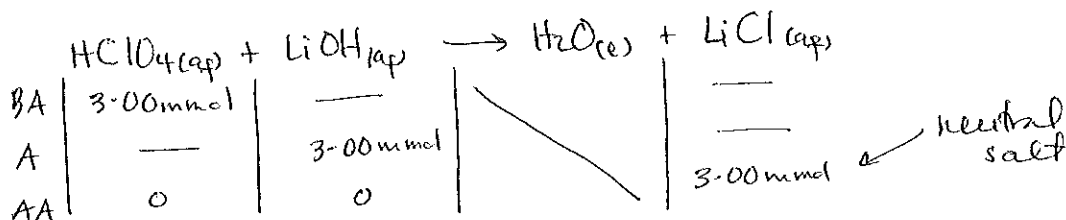
$$[\text{H}_3\text{O}^+] = [\text{HClO}_4] = 0.150M \quad \text{pH} = -\log [\text{H}_3\text{O}^+] = -\log(0.150) = \underline{0.824}$$

(b) Excess HClO_4
(10.00 mL LiOH)



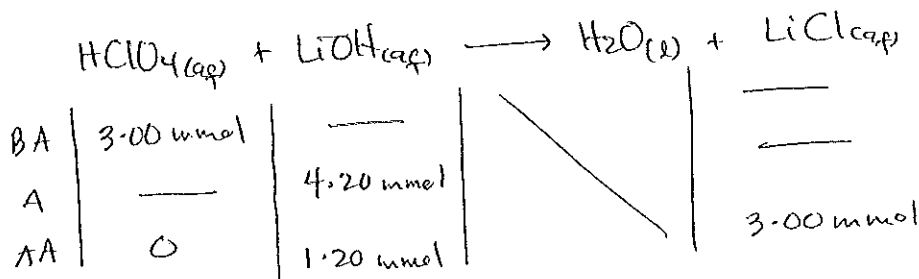
$$[\text{H}_3\text{O}^+] = [\text{HClO}_4] = \frac{1.80 \text{ mmol}}{30.0 \text{ mL}} = 0.0600M \quad \text{pH} = -\log(0.0600) = \underline{1.222}$$

(c) Equivalence Point
(25.00 mL LiOH)



$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7}M \quad \text{pH} = -\log(1.0 \times 10^{-7}) = \underline{7.00}$$

(d) Excess LiOH
(35.00 mL LiOH)



$$[\text{OH}^-] = [\text{LiOH}]_{\text{excess}} = \frac{1.20 \text{ mmol}}{55.0 \text{ mL}} = 0.0218M$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log(0.0218) = 1.661$$

$$\text{pH} = 14.00 - 1.661 = \underline{12.34}$$

